Chemical Analysis & Heat Transfer Answer Key

CHY 113 | Guided group activity |

Work together; do not divide the questions. Explain your reasoning before calculating. Predict signs and check units. The activity is intentionally longer than one class period; unfinished questions can be completed independently.

Model 1 — Solution Stoichiometry: Start with the Chemistry

Two aqueous solutions are mixed: 150. mL of 0.500 M $AgNO_3$ and 100. mL of 0.400 M K_2CrO_4.

$2\ AgNO_3(aq) + K_2CrO_4(aq) \rightarrow Ag_2CrO_4(s) + 2\ KNO_3(aq)$

1. Without calculating, explain why you cannot determine the precipitate mass from the concentration of just one reactant.

Need BOTH reactant amounts: either may run out first. Precipitate is controlled by the limiting reactant.

2. Determine the limiting reactant. Show the mole comparison using the balanced equation.

$AgNO_3$: 0.150 L × 0.500 M = 0.0750 mol

K_2CrO_4: 0.100 L × 0.400 M = 0.0400 mol

0.0750/2 = 0.0375 mol Ag_2CrO_4 vs. 0.0400/1 = 0.0400 mol → $AgNO_3$ limits.

3. Calculate the theoretical mass of Ag_2CrO_4. Include units and a reasonable number of significant figures.

0.0375 mol Ag_2CrO_4 × 331.73 g/mol = 12.4 g Ag_2CrO_4 ✓

4. If additional K_2CrO_4 were added after the reaction, would more precipitate necessarily form? Explain.

No. All Ag^+ has been consumed ($AgNO_3$ limiting). Extra chromate alone cannot make more Ag_2CrO_4.

Model 2 — Combustion Analysis: Working Backward

A 0.500 g sample of a compound containing only C, H, and O is completely burned. It produces 0.733 g CO_2 and 0.300 g H_2O.

5. Why can the carbon and hydrogen amounts be obtained directly from the combustion products, but not the oxygen amount?

All C ends up in CO_2 and all H in H_2O. O in the products also comes from added O_2.

6. Determine moles of C and moles of H in the original sample. Show both conversions, including the 2:1 H:H_2O ratio.

n(C) = 0.733 g CO_2 ÷ 44.01 = 0.0167 mol C

n(H) = (0.300 g H_2O ÷ 18.02) × 2 = 0.0333 mol H

7. Determine the mass and moles of oxygen originally present. Why must the original sample contain only C, H, and O for this subtraction to work?

m(C) ≈ 0.200 g; m(H) ≈ 0.0336 g

m(O) = 0.500 − 0.200 − 0.0336 = 0.266 g

n(O) = 0.266/16.00 ≈ 0.0166 mol. Subtraction works only if NO other elements are present.

8. Determine the empirical formula. What is the final step if your mole ratios are not close to whole numbers?

C : H : O ≈ 0.0167 : 0.0333 : 0.0166 → 1 : 2 : 1 → CH_2O.

If fractional ratios remain, multiply ALL ratios by the smallest whole number that clears them.

9. The molar mass of this compound is approximately 180 g/mol. Determine its molecular formula. Explain how you use the empirical-formula mass.

M(CH_2O) ≈ 30.03 g/mol; 180/30.03 ≈ 6.

(CH_2O)$_6$ = $C_6H_{12}O_6$ ✓

Model 3 — Heat, Thermal Energy, and Direction

Heat (q) is energy transferred because of a temperature difference. The sign of q is defined for the system being considered; q is not the total thermal energy stored in an object.

10. A hot metal block is placed into cooler water. Identify the direction of heat transfer and predict the sign of q for each substance.

Heat flows hot metal → cooler water. q(metal) < 0; q(water) > 0.

11. The metal and water eventually reach the same temperature. Does that mean they contain the same total thermal energy? Explain.

No! Same temperature means equal average particle kinetic energy, not equal total thermal energy. Mass and material matter.

12. An instant cold pack becomes cold as a solid dissolves. Treat the dissolving contents as the system. Is the process endothermic or exothermic? What is the sign of q for the system? Where does the energy come from?

Endothermic dissolution: q(system) > 0. Energy enters the dissolving contents from the surrounding pack/hand/air.

13. Concentrated acid is carefully added to water and the mixture warms. Treat the dilution/mixing process as the system. Is the process endothermic or exothermic? What is the sign of q for that process? How can the solution temperature nevertheless increase?

Exothermic mixing/dilution: q(process) < 0. The released energy warms the solution (q for the warming solution > 0).

Always specify WHICH system is being described!

14. A student writes: “The metal has negative heat because it is hot.” Explain what is wrong with this statement.

An object does not “have negative heat.” q describes energy TRANSFER, not stored heat. Hot metal cooling loses energy → q < 0.

Model 4 — Using q = mcΔT

For a substance undergoing a temperature change without a phase change: q = mcΔT, where ΔT = Tfinal − Tinitial.

15. A 50.0 g sample of water warms from 25.0°C to 33.4°C. Predict the sign of q, then calculate q. Use cwater = 4.184 J/(g·°C).

Water warms → q > 0. ΔT = 33.4 - 25.0 = +8.4°C

q = (50.0 g)(4.184 J/g°C)(+8.4°C) = +1.8 × 10^3 J (2 sig figs).

16. The same 50.0 g sample cools from 33.4°C to 25.0°C. Without repeating every calculation, predict q and explain the relationship to Question 15.

Same size ΔT, opposite sign → q = -1.8 × 10^3 J. Water loses the energy it gained in #15.

17. A 25.0 g metal sample absorbs 2.33 kJ and its temperature increases by 200.0°C. Calculate its specific heat. What unit conversions are needed?

2.33 kJ = 2330 J. c = q/(mΔT) = 2330/(25.0 × 200.0)

= 0.466 J/(g°C).

18. Equal masses of water and copper absorb equal amounts of heat, with no phase changes. Which undergoes the larger temperature increase? Explain using specific heat (cwater = 4.184; ccopper = 0.385 J/(g·°C)).

Copper has the larger ΔT. For equal q and m, ΔT = q/(mc); lower c means greater temperature change.

19. A burning match and a bonfire can have similar temperatures. Why is the bonfire much more effective at warming nearby people? Use the distinction between temperature and energy transfer.

Temperature is not total energy. The bonfire has far more burning material and can transfer much more energy per unit time.

Model 5 — Calorimetry: Two Substances, One Energy Transfer

A 28.7 g metal sample at 99.8°C is placed in 225 g of water initially at 21.0°C. The final temperature is 23.1°C. Assume no heat is exchanged with the container or surroundings.

20. Before doing arithmetic, explain why the final temperature must fall between the two initial temperatures.

With no heat loss, hot metal cools and cold water warms until they meet at a temperature between 21.0°C and 99.8°C.

21. Calculate ΔTwater and ΔTmetal. Include their signs.

ΔT(water) = 23.1 − 21.0 = +2.1°C

ΔT(metal) = 23.1 − 99.8 = −76.7°C

22. Calculate qwater, then determine qmetal using conservation of energy. Explain why the signs differ.

q(water) = (225)(4.184)(+2.1) = +1.98 × 10^3 J

q(metal) = −q(water) = −1.98 × 10^3 J. Equal magnitudes; opposite directions.

23. Solve for the specific heat of the metal. Compare your result with aluminum, c = 0.897 J/(g·°C). Report appropriate significant figures.

c(metal) = −1.98 × 10^3 / [(28.7)(−76.7)] ≈ 0.90 J/(g°C) (2 sig figs).

Very close to aluminum (0.897).

24. A student instead writes qmetal = qwater and obtains a negative specific heat. Diagnose the error using the physical direction of energy flow.

They omitted the negative sign in q(metal) = −q(water). Metal cools → q(metal) < 0; its c must be positive.

Model 6 — Transfer and Check Your Reasoning

A 59.047 g metal sample is heated to 100.0°C and placed in 100.0 g of water initially at 23.7°C. The final temperature is 27.8°C. Assume the water and metal exchange heat only with each other.

25. Create a plan for determining the metal's specific heat. List the quantities you will calculate and the sign you expect for each heat transfer.

Plan: ΔT(water) positive; q(water) positive; q(metal) = −q(water) negative; ΔT(metal) negative; c = q/(mΔT).

26. Calculate the specific heat of the metal. Show the complete setup with units and appropriate significant figures.

ΔT(water) = +4.1°C; q(water) = 100.0 × 4.184 × 4.1 = +1715 J

ΔT(metal) = 27.8 − 100.0 = −72.2°C

c = −1715/[59.047 × (−72.2)] ≈ 0.40 J/(g°C) (2 sig figs).

27. Suppose the cup also absorbed some heat, but you ignored it. Would your calculated metal specific heat be too high or too low? Explain.

Too LOW. Some metal energy heats the cup, so |q(metal)| > q(water). Ignoring cup heat undercounts energy lost.

Final reflection — A reliable thermochemistry strategy

28. Write a three-step checklist you can use to catch sign errors before finishing any thermochemistry calculation.

1) Define system and predict heat-flow direction.

2) Calculate ΔT = Tfinal − Tinitial; check sign of q.

3) For two bodies, enforce q(lost) + q(gained) = 0; verify units and physical sense.

Independent Practice — Apply the Methods Without Scaffolding

Show your setup, units, and final answer. These problems are intentionally less guided than the models above.

29. What volume of 0.496 M HCl is needed to neutralize 20.0 mL of 0.809 M $Sr(OH)_2$? Write a balanced equation.

$Sr(OH)_2 + 2\ HCl \rightarrow SrCl_2 + 2\ H_2O$

$n[Sr(OH)_2] = 0.0200 \times 0.809 = 0.01618$ mol

$n(HCl) = 2 \times 0.01618 = 0.03236$ mol; $V = n/M = 0.0652$ L = 65.2 mL.

30. A 25.0 mL HCl sample requires 15.0 mL of 0.200 M NaOH to reach the equivalence point. Find the molarity of HCl.

$HCl + NaOH \rightarrow NaCl + H_2O$ (1:1)

$n(NaOH) = 0.0150\ L \times 0.200\ M = 0.00300$ mol

$M(HCl) = 0.00300/0.0250\ L = 0.120\ M$.

31. A 0.105 g compound containing only C, H, and O produces 0.257 g CO_2 and 0.0350 g H_2O upon combustion. Find its empirical formula. If its molar mass is approximately 108 g/mol, find its molecular formula.

$n(C) = 0.257/44.01 = 0.00584$; $n(H) = 2(0.0350/18.02) = 0.00388$

$m(O) = 0.105 - m(C) - m(H) \approx 0.0309$ g $\rightarrow n(O) \approx 0.00193$

C:H:O ≈ 3:2:1 → C_3H_2O; 108/54.05 ≈ 2 → $C_6H_4O_2$.

32. How much heat is needed to warm 7.89 g of ethanol from 25.0°C to 72.7°C? $c_{ethanol}$ = 2.419 J/(g·°C).

$\Delta T = 72.7 - 25.0 = +47.7$°C

$q = (7.89)(2.419)(47.7) = +911$ J.

33. A 500.0 g piece of lead cools from 75.0°C to 25.0°C. Calculate q_{lead}. c_{lead} = 0.129 J/(g·°C). State whether the lead gains or loses energy.

$\Delta T = 25.0 - 75.0 = -50.0$°C

$q = (500.0)(0.129)(-50.0) = -3.23 \times 10^3$ J = −3.23 kJ. Lead loses energy.

34. A 1.00 kg piece of copper cools to 100.0°C and releases 231 kJ. What was its initial temperature? c_{copper} = 0.385 J/(g·°C).

$q = -231{,}000$ J; $\Delta T = q/(mc) = -231000/(1000 \times 0.385) = -600$°C

100.0 − T(initial) = −600 → T(initial) = 700°C (approx.).

35. A 44.7 g unknown substance absorbs 2110 J as it warms from 23.2°C to 89.6°C. Calculate its specific heat.

$\Delta T = 89.6 - 23.2 = 66.4$°C

c = 2110/[44.7 × 66.4] = 0.711 J/(g°C).

36. A 360.0 g steel rod is placed in 425 g of water initially at 24.0°C. The final temperature is 42.7°C. If csteel = 0.449 J/(g·°C), determine the rod's initial temperature. Assume no heat is exchanged with the surroundings.

q(water) = (425)(4.184)(42.7 − 24.0) ≈ +33,250 J

q(steel) = −33,250 J = (360.0)(0.449)(42.7 − T(initial))

T(initial) ≈ 248°C.

37. A 75.0 g metal sample at 95.0°C is placed into 150.0 g water at 20.0°C. They reach 24.0°C. Determine the specific heat of the metal, assuming no heat loss to the cup or air.

q(water) = (150.0)(4.184)(24.0 − 20.0) = +2510 J

ΔT(metal) = 24.0 − 95.0 = −71.0°C

c(metal) = −2510/[75.0 × (−71.0)] ≈ 0.47 J/(g°C).